

Chapter 8: Rates and equilibrium 1

Knowledge organiser

Rates of reaction

The **rate of a reaction** is how quickly the reactants turn into the products.

To calculate the rate of a reaction, you can measure:

- how quickly a reactant is used up

$$\text{mean rate of reaction} = \frac{\text{quantity of reactant used}}{\text{time taken}}$$

- how quickly a product is produced.

$$\text{mean rate of reaction} = \frac{\text{quantity of product formed}}{\text{time taken}}$$

For reactions that involve a gas, this can be done by measuring how the mass of the reaction changes or the volume of gas given off by the reaction.

Volume of gas produced

The reaction mixture is connected to a gas syringe or an upside down measuring cylinder. As the reaction proceeds the gas is collected.

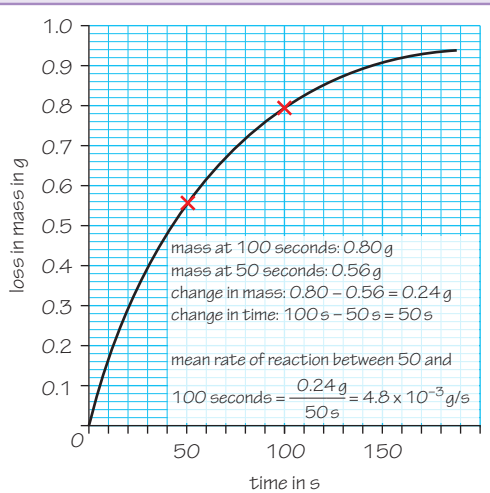
The rate for the reaction is then:

$$\text{rate} = \frac{\text{volume of gas produced}}{\text{time taken}}$$

Volume is measured in cm^3 and time in seconds, so the unit for rate is cm^3/s .

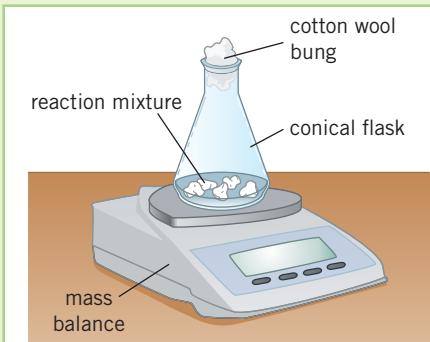
Mean rate between two points in time

To get the mean rate of reaction between two points in time:



Change in mass

The reaction mixture is placed on a mass balance. As the reaction proceeds and the gaseous product is given off, the mass of the flask will decrease.



The rate for the reaction is then:

$$\text{rate} = \frac{\text{change in the mass}}{\text{time taken}}$$

The mass is measured in grams and time is measured in seconds. Therefore, the unit of rate is g/s .

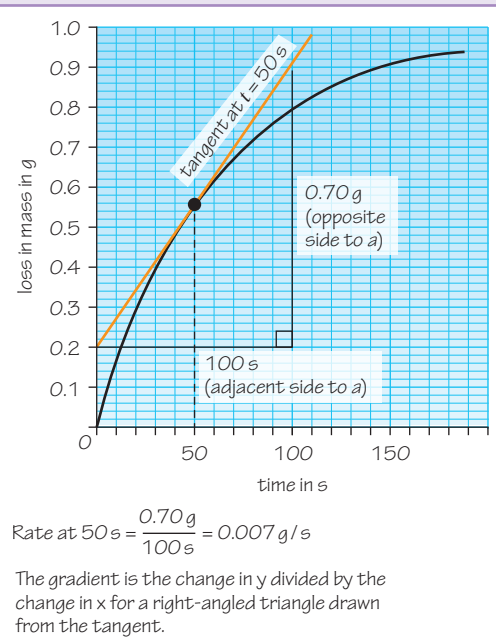
Calculating rate from graphs (HT only)

The results from an experiment can be plotted on a graph.

- A steep gradient means a high rate of reaction – the reaction happens quickly.
- A shallow gradient means a low rate of reaction – the reaction happens slowly.

Mean rate at specific time

To obtain the rate at a specific time draw a **tangent** to the graph and calculate its **gradient**.



Collision theory

For a reaction to occur, the reactant particles need to collide. When the particles collide, they need to have enough energy to react or they will just bounce apart. This amount of energy is called the **activation energy**.

You can increase the rate of a reaction by:

- increasing the **frequency of collisions**
- increasing the energy of the particles when they collide.

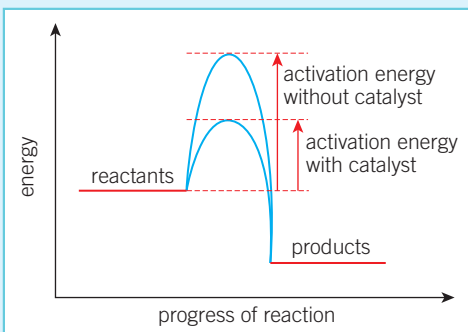
Factors affecting rate of reaction

Condition that increases rate	How is this condition caused?	Why it has that effect
increasing the temperature	Heat the container in which the reaction is taking place.	1 particles move faster, leading to more frequent collisions 2 particles have more energy, so more collisions result in a reaction note that these are two <i>separate</i> effects
increasing the concentration of solutions	Use a solution with more solute in the same volume of solvent.	there are more reactant particles in the reaction mixture, so collisions become more frequent
increasing the pressure of gases	Increase the number of gas particles you have in the container or make the container smaller.	less space between particles means more frequent collisions
increasing the surface area of solids	Cut the solid into smaller pieces, or grind it to create a powder, increasing the surface area. Larger pieces decrease the surface area.	only reactant particles on the surface of a solid are able to collide and react; the greater the surface area the more reactant particles are exposed, leading to more frequent collisions

Catalysts

Some reactions have specific substances called **catalysts** that can be added to increase the rate. These substances are not used up in the reaction.

A catalyst provides a different reaction pathway that has a lower activation energy. As such, more particles will collide with enough energy to react, so more collisions result in a reaction.



Chapter 8: Rates and equilibrium 2

Knowledge organiser

Reaction conditions

The conditions of a reaction refer to the external environment of the reaction. When the reaction occurs in a closed system, you can change the conditions by:

- changing the concentration of one of the substances
- changing the temperature of the entire reaction vessel
- changing the pressure inside the vessel.

Le Châtelier's principle (HT only)

At equilibrium, the amount of reactants and products is constant. In order to change the amounts of reactant and product at equilibrium the *conditions* of the reaction must be changed. The closed system will then counteract the change by favouring either the forward reaction or the reverse reaction. This is known as **Le Châtelier's principle**. For example, lowering the concentration of the product in the system causes the forward reaction to be **favoured** to increase the concentration of the product.

Changing concentrations (HT only)

Change	Effect	Explanation
decrease concentration of product	favours the forward reaction	opposes the change by making <i>less</i> reactant and <i>more</i> product
increase concentration of product	favours the reverse reaction	opposes the change by making <i>more</i> reactant and <i>less</i> product
decrease concentration of reactant	favours the reverse reaction	opposes the change by making <i>more</i> reactant and <i>less</i> product
increase concentration of reactant	favours the forward reaction	opposes the change by making <i>less</i> reactant and <i>more</i> product

Changing temperature (HT only)

Change	Effect	Explanation
increase temperature of surroundings	favours the endothermic reaction	opposes the change by decreasing the temperature of the surroundings
decrease temperature of surroundings	favours the exothermic reaction	opposes the change by increasing the temperature of the surroundings

Changing pressure (HT only)

Change	Effect	Explanation
increase the pressure	favours the reaction that results in fewer molecules	decreasing the number of molecules within the vessel opposes the change because it decrease pressure
decrease the pressure	favours the direction that results in more molecules	increasing the number of molecules within the vessel opposes the change because it increase pressure



Key terms

Make sure you can write a definition for these key terms.

activation energy catalyst collision collision theory closed system
conditions dynamic equilibrium frequency of collision gradient
Le Châtelier's principle rate of reaction reversible reaction tangent

Reversible reactions

In some reactions, the products can react to produce the original reactants. This is called a **reversible reaction**. When writing chemical equations for reversible reactions, use the \rightleftharpoons symbol.

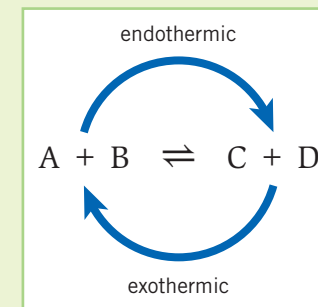
In this reaction:

- A and B can react to form C and D – the forward reaction
- C and D can react to form A and B – the reverse reaction.

The different directions of the reaction have opposite energy changes.

If the forward reaction is *endothermic*, the reverse reaction will be *exothermic*.

The same amount of energy is transferred in each direction.



Equilibrium

In a **closed system** no reactants or products can escape. If a reversible reaction is carried out in a closed system, it will eventually reach **dynamic equilibrium** – a point in time when the forward and reverse reactions have the same rate.

At dynamic equilibrium:

- the reactants are still turning into the products
- the products are still turning back into the reactants
- *the rates* of these two processes are *equal*, so overall the amount of reactants and products are constant.

Dynamic equilibrium

At dynamic equilibrium the amount of reactant and product are constant, but not necessarily equal.

You could have a mixture of reactants and products in a 50:50 ratio, in a 75:25 ratio, or in any ratio at all. The **conditions** of the reaction are what change that ratio.

How dynamic equilibrium is reached

Progress of reaction	start of reaction	middle of reaction	at dynamic equilibrium
Amount of A + B	high	decreasing	constant
Frequency of collisions A + B	high	decreasing	constant
Rate of forward reaction	high	decreasing	same as rate of reverse reaction
<p>rate of reaction</p> <p>time</p> <p>forward reaction</p> <p>reverse reaction</p> <p>equilibrium is reached at this point</p>			
Amount of C + D	zero	increasing	constant
Frequency of collisions C + D	no collisions	increasing	constant
Rate of reverse reaction	zero	increasing	same as rate of forward reaction